

Chapter	1 Periodic Table and Periodic Properties	Class	1 st Year	Subject	Chemistry
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Give Short Answers of the Following Questions

1. How are elements arranged in the modern periodic table?

Ans. Elements in the modern periodic table are arranged in order of increasing atomic number. Elements with similar chemical properties are placed in the same group, and the pattern of properties repeats periodically.

2. What do the s, p, d, and f blocks in the periodic table represent?

Ans. These blocks (s, p, d, and f) represent the type of atomic orbital being filled by the valence electrons. The s-block includes groups 1 and 2, the p-block includes group 13 to 18, the d-block contains transition metals and the f-block includes lanthanides and actinides.

3. What does the periodic number of an element indicate?

Ans. The period number indicates the number of occupied electron shells in an atom of the element. For example, an element in period 3 has electrons in three energy levels.

4. We determine the electron configuration of an element using its position on the periodic table. How?

Ans. The element of the period number represents the highest principal energy level, and the block shows the type of orbital. The group number of the element helps to determine the number of electrons in the outermost shell.

5. How can we identify an unknown element based on its physical and chemical properties?

Ans. We can identify an unknown element based on its physical and chemical properties by comparing the properties of an elements such as reactivity, ion formation, and bonding behavior with known trends I the periodic table, we can estimate its possible group and period and identify it.

6. How can we predict the properties of elements in the same group?

Ans. Elements in the same group have the same number of valence electrons, so they exhibit similar chemical properties and reactions. For example, group 1 elements all react with water to form hydroxides and release hydrogen gas.

7. How do atomic radius and ionization energy change across a period and down a group?

Ans. Across a period: The atomic radius decreases and ionization energy increases across a period due to greater nuclear attraction on the valence shell electrons.

Down a group: The atomic radius increases and ionization energy down the group decreases due to added electron shells and increases shielding. So nuclear attractions decreases.

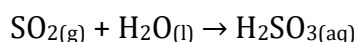
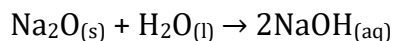
8. How does the number of valence electrons affect metallic and nonmetallic character?

Ans. Elements with fewer valence electrons (like 1 to 3) tend to lose them easily and show metallic character, on the other hand, those elements with more valence electrons (like 5 to 7) gain electrons and show nonmetallic character.

9. How do oxides of metals and non-metals behave in water?

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Ans. The metal oxides usually react with water to form basic solutions, on the other hand, non-metal oxides form acidic solutions. For example, Na₂O forms NaOH, and SO₃ from H₂SO₄.



10. How does the reactivity of group 1 and group 17 elements change down the group?

Ans. In group 1, reactivity of elements increases down the group because the outermost electron is easily lost. In group 17, reactivity of elements decreases down the group as gaining an electron becomes harder.

11. How do transition metals differ from s-block elements?

Ans. The transition metals differ from s-block elements because of following aspects:

- Transition metals can form multiple oxidation states.
- They often form colored compounds.
- They act as catalysts.
- They involve the filling of d-orbitals.

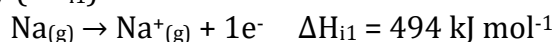
12. Why do non-metals tend to gain electrons while metals tend to lose them?

Ans. The non-metals have high Electronegativity and nearly full valence shells, so that is why, they tend to gain electrons. On the other hand, metals have low ionization energy and few valence electrons, that is why, they lose them easily.

13. What is 1st ionization energy? Give an example.

Ans. 1st ionization energy:

The energy needed to remove one electron from each atom in one mole of atoms of the element in the gaseous state to form one mole of gaseous 1⁺ ions is known as 1st ionization energy (ΔH_{i1})."



14. Explain why sulphur has lower first ionization energy than phosphorus.

Ans. Sulphur has lower first ionization energy than phosphorus due to electron-electron repulsion in its 3p orbitals, which makes it easier to remove an electron.

Electronic configuration:

- Phosphorus: 1s² 2s² 2p⁶ 3s² 3p³
- Sulphur: 1s² 2s² 2p⁶ 3s² 3p⁴

Half-filled stability (phosphorus):

- Phosphorus has a half-filled 3p subshell (3p³), which is relatively stable due to symmetrical distribution and minimized electron repulsion.
- This stability makes it harder to remove one of its electrons.

Increased electron repulsion in sulphur:

- Sulphur has one more electron than phosphorus, giving it a 3p⁴ configuration.
- In this case, one of the 3p orbitals contains a pair of electrons, which experience increased repulsion.

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- This repulsion makes it easier to remove one of the paired electrons, hence lowering the ionization energy.

15. Why the elements in Group 13 to 17 are called p-block elements?

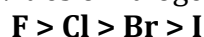
Ans. In group 13 to 17 elements, the last electron enters in p-orbital of their respective shells. That is why they are called p-block elements e.g., B, C, N, O, F.

16. What are the factors that affect Electronegativity?

Ans. Factors Affecting Electronegativity

i) Atomic size

A large atomic size will result in a lower value of Electronegativity. This is because electrons being far away from the nucleus will experience a weaker force of attraction. For example, the electronegativities of halogens in group 17 are in the order:



ii) Effective nuclear charge

A higher value of the effective nuclear charge will result in a greater value of electronegativity, because an increase in nuclear charge causes greater attraction to the bonded electrons. This is why the Electronegativity in a period increases from left to right. The electronegativity of Li in period 2 is 1.0 and F has a value of 4.0.

17. What factors are responsible for the increasing reactivity of alkali metals as you move down the group?

Ans. Down the group (in alkali metals), atomic size increases and ionization energy decreases. It results in greater tendency to form positive ions and they react more readily.

18. Why some of the elements show variable oxidation number while others do not?

Ans. The variability in oxidation numbers among elements is primarily due to difference in electronic configurations, atomic structure, and position in the periodic table. Here's a breakdown of why some elements show variable oxidation states while others do not.

1. Transition Metals and Variable States

- Transition metals (like iron, copper, manganese) show variable oxidation states because they have partially filled d-orbitals.
- Their outer electron configuration typically ends in $(n-1) d^{1-10}ns^{0-2}$, so both s-and d-electrons can be involved in bonding.
- Example:** Iron (Fe) can lose 2 electrons to become Fe^{3+} .

2. Relativistic and Inert pair Effect (in heavier elements)

- Heavier p-block elements (like Sn, Pb) show variable oxidation states due to the inert pair effect, where the s-electrons are less available for bonding due to relativistic stabilization.
- Example:** Lead (Pb) shows both +2 and +4 oxidation states.

3. Main Group Elements and Fixed Oxidation States

- Group 1 and group 2 metals (like Na, Mg) almost always show a fixed oxidation number (+1 and +2 respectively) because they lose their valence s-electrons easily, and further loss would require breaking into a stable inner shell, which is energetically unfavorable.
- Similarly, nonmetals like fluorine always have an oxidation number of -1 because it is the most electronegative element and can only gain one electron to complete its octet.

4. Stability of Electronic Configurations

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- Some elements achieve stable noble gas configurations after gaining or losing a specific number of electrons. Once this stability is reached, further changes are energetically unfavorable.
- For example, oxygen typically forms -2 oxidation state to complete its octet.

19. Identify the element which is in period 5 and group 15?

Ans. Antimony (Sb)

20. Why oxides of sodium and magnesium are more ionic than the oxides of nitrogen and phosphorous?

Ans. Oxides of sodium (Na) and magnesium (Mg) are more ionic than the oxides of nitrogen (N) and phosphorus (P) due to differences in Electronegativity, metallic character, and bonding tendencies of these elements.

1. Nature of Elements (Metals and Non-metals)

- Sodium and magnesium are metals (from groups 1 and 2). Metals tend to lose electrons easily and form cations.
- Nitrogen and phosphorus are non-metals (Group 15). Non-metals prefer to share electrons rather than form ions, leading to covalent bonding.

2. Electronegativity Difference

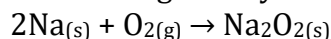
- Ionic character is stronger when there is a large Electronegativity difference between the metal and oxygen.
- Sodium (0.93) and magnesium (1.31) have much lower electronegativities than oxygen (3.44), covalent bonds.
- Nitrogen (3.04) and phosphorus (2.19) have closer Electronegativities values to oxygen, leading covalent bonds.

3. Bonds Formation Tendencies

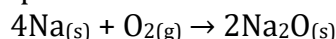
- Na₂O and MgO form ionic lattices, with strong electrostatic attraction between ions (e.g., Na⁺ and O²⁻).
- N₂O₅ and P₄O₁₀, oxides of nitrogen and phosphorus, exist as molecular compounds with covalent bonds (electron sharing).

21. Give reason for the different chemical reactivities of Na and Mg toward oxygen and chlorine.

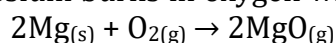
Ans. Sodium burns in oxygen with a golden yellow flame to produce a white solid mixture of sodium oxide and sodium peroxide. Sodium is kept under kerosene oil to prevent its reaction with air. It reacts vigorously with oxygen in open air to form sodium peroxide.



Under special conditions like limited O₂ or high temperature, sodium oxide is formed.

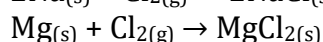
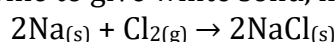


Magnesium burns in oxygen with an intense white flame to give white solid magnesium oxide.



With Chlorine

Chlorine reacts with both metals to give soluble salts. It reacts exothermically with sodium, golden yellow flame is seen and white solid, sodium chloride is formed. Magnesium also reacts with chlorine to give white solid, magnesium chloride.



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22. Why the ionization energy of lithium is much lower than that of helium despite the fact that the nuclear charge of lithium is +3 and that of helium is +2.

Ans. The ionization energy of lithium is much lower than that of helium because of the following key factors:

1. Electron Configuration and Shell Structure

Helium has the electron configuration: $1s^2$

Both electrons are in the first energy level, very close to the nucleus and strongly held.

Lithium has the electron configuration: $1s^2 2s^1$

The outermost electron ($2s$) is in the second energy level, which is farther from the nucleus and shielded by the inner $1s$ electrons.

2. Electron Shielding Effect

In lithium, the two inner $1s$ electrons shield the outer $2s$ electron from the full +3 nuclear charge.

As a result, the effective nuclear charge felt by the $2s$ electron is significantly less than +3, it is roughly around +1.

In helium, there's no shielding, both electrons are strongly attracted to the nucleus.

3. Distance from Nucleus

The $2s$ electron in lithium is farther from the nucleus compared to helium's $1s$ electrons.

Greater distance = weaker attraction = easier to remove = lower ionization energy.

23. The ionization energy of Be (atomic no. 4) is higher than that of B (atomic no. 5), despite the fact that the nuclear charge of Be is +4 and that of B is +5.

Ans. Beryllium (Be) has a higher ionization energy than boron (B) despite having a lower nuclear charge.

Beryllium (Be): Atomic number = 4, Electronic configuration: $1s^2 2s^2$

Boron (B): Atomic number = 5, Electronic configuration: $1s^2 2s^2 2p^1$

- In Be, the electron is removed from a $2s$ orbital, which is lower in energy, more penetrating, and closer to the nucleus.
- In B, the electron is removed from a $2p$ orbital, which is Higher in energy, farther from the nucleus, less penetrating and more shielded by the inner $1s$ and $2s$ electrons. So, even though B has a higher nuclear charge (+5 vs. +4), the $2p$ electron is easier to remove than the $2s$ electron in Be.
- So, $2p$ orbitals are less tightly bound than $2s$ orbital, despite being in the same principal energy level ($n = 2$).

This makes the first ionization energy of Be lower, because you're removing an electron that's already less stable.

24. What is common in Na^+ , Mg^{2+} , Al^{3+} , Ne^0 and F^- ? Arrange them in increasing order of sizes.

Ans. $Al^{3+} < Mg^{2+} < Na^+ < Ne^0 < F^-$

They all have 10 electrons each so they are iso-electronic species.

As the nuclear charge increases (with constant number of electrons), the electrons cloud is pulled in tighter, and the radius decreases.

Al^{3+} : Higher nuclear charge with 10 electrons → most tightly held → smallest

F^- : Lowest nuclear charge with 10 electrons → weakest pull on electrons → largest

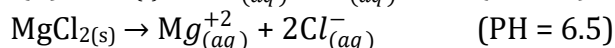
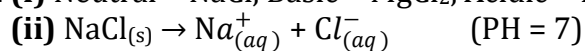
25. Consider the chlorides of sodium, magnesium, and phosphorus (v): $NaCl$, $MgCl_2$, and PCl_5 .

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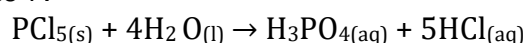
(i) Classify each of these chlorides as acidic, basic, or neutral.

(ii) For each chloride, briefly explain the reason for your classification, referring to their behavior when dissolved in water.

Ans. (i) Neutral = NaCl, Basic = MgCl₂, Acidic = PCl₅



- Neutral chlorides are salts that, when dissolved in water, produce a neutral solution with a pH close to 7.



- PCl₅ reacts with water to produce phosphoric acid and hydrochloric acid i.e. produces acidic solution.